

# Water and Life

# 3



▲ **Figure 3.1** How does life on Earth depend on the chemistry of water?

## KEY CONCEPTS

- 3.1** Polar covalent bonds in water molecules result in hydrogen bonding
- 3.2** Four emergent properties of water contribute to Earth's suitability for life
- 3.3** Acidic and basic conditions affect living organisms



▲ Black guillemots, threatened by climate change

## The Molecule That Supports All of Life

Life on Earth began in water and evolved there for 3 billion years before spreading onto land. Water is the substance that makes life possible as we know it here on Earth, and possibly on other planets as well. All organisms familiar to us are made mostly of water and live in an environment dominated by water.

Three-quarters of Earth's surface is covered by water. Although most of this water is in liquid form, water is also present on Earth as a solid (ice) and a gas (water vapor). Water is the only common substance on Earth to exist in the natural environment in all three physical states of matter. Furthermore, the solid form of water floats on the liquid form, a rare property emerging from the chemistry of the water molecule.

As the Earth is warming from climate change (see Concept 1.1), the ratio of ice to liquid water is changing. Arctic sea ice and glaciers are melting, affecting life on, under, and around them. In the Arctic, warmer waters and the smaller ice pack are resulting in blooms of phytoplankton (microscopic aquatic photosynthetic organisms), seen from space as the “cloudy” seawater in **Figure 3.1**. Organisms that depend on Arctic ice, however, are suffering. For instance, a population of black guillemots in Alaska is declining due to the warming climate and reduction of Arctic sea ice.

In this chapter, you will learn how the structure of a water molecule allows it to interact with other molecules, including other water molecules. This ability leads to water's unique emergent properties that help make Earth suitable for life.

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Get Ready for This Chapter

## CONCEPT 3.1

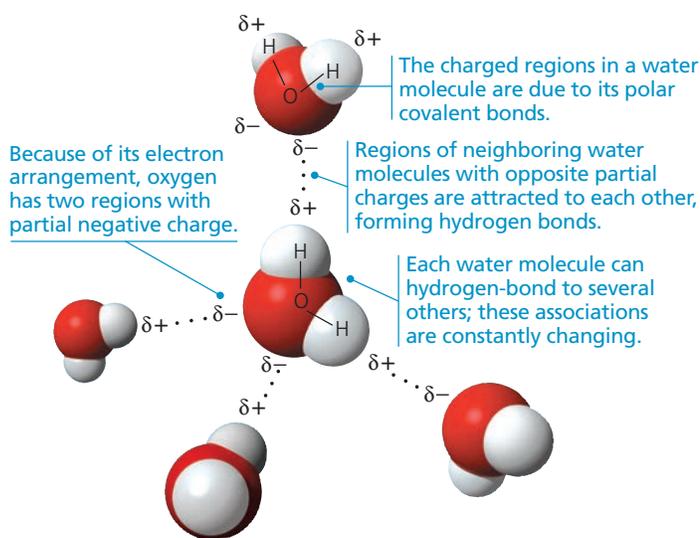
### Polar covalent bonds in water molecules result in hydrogen bonding

Water is so familiar to us that it is easy to overlook its many extraordinary qualities. Following the theme of emergent properties, we can trace water's unique behavior to the structure and interactions of its molecules.

Studied on its own, the water molecule is deceptively simple. It is shaped like a wide V, with its two hydrogen atoms joined to the oxygen atom by single covalent bonds. Oxygen is more electronegative than hydrogen, so the electrons of the covalent bonds spend more time closer to oxygen than to hydrogen; these are **polar covalent bonds** (see Figure 2.11). This unequal sharing of electrons and water's V-like shape make it a **polar molecule**, meaning that its overall charge is unevenly distributed. In water, the oxygen of the molecule has two regions of partial negative charge ( $\delta^-$ ), and each hydrogen has a partial positive charge ( $\delta^+$ ).

The properties of water arise from attractions between oppositely charged atoms of different water molecules: The partially positive hydrogen of one molecule is attracted to the partially negative oxygen of a nearby molecule. The two molecules are thus held together by a hydrogen bond (Figure 3.2). When water is in its liquid form, its hydrogen bonds are very fragile, each only about 1/20 as strong as a covalent bond. The hydrogen bonds form, break, and re-form with great frequency. Each lasts only a few trillionths of a second, but the molecules are constantly forming new hydrogen bonds with a succession of partners. Therefore, at any instant, most of the water molecules

▼ **Figure 3.2** Hydrogen bonds between water molecules.



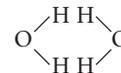
**DRAW IT** ► Draw partial charges on the water molecule at the far left, and draw three more water molecules hydrogen-bonded to it.

**Animation: Polarity of Water**

are hydrogen-bonded to their neighbors. The extraordinary properties of water emerge from this hydrogen bonding, which organizes water molecules into a higher level of structural order.

## CONCEPT CHECK 3.1

- 1. MAKE CONNECTIONS** ► What is electronegativity, and how does it affect interactions between water molecules? (Review Figure 2.11.)
- 2. VISUAL SKILLS** ► Look at Figure 3.2 and explain why the central water molecule can hydrogen bond to four (rather than three or five) other water molecules.
- 3.** Why is it unlikely that two neighboring water molecules would be arranged like this?



- 4. WHAT IF?** ► What would be the effect on the properties of the water molecule if oxygen and hydrogen had equal electronegativity?

For suggested answers, see Appendix A.

## CONCEPT 3.2

### Four emergent properties of water contribute to Earth's suitability for life

We will examine four emergent properties of water that contribute to Earth's suitability as an environment for life: cohesive behavior, ability to moderate temperature, expansion upon freezing, and versatility as a solvent.

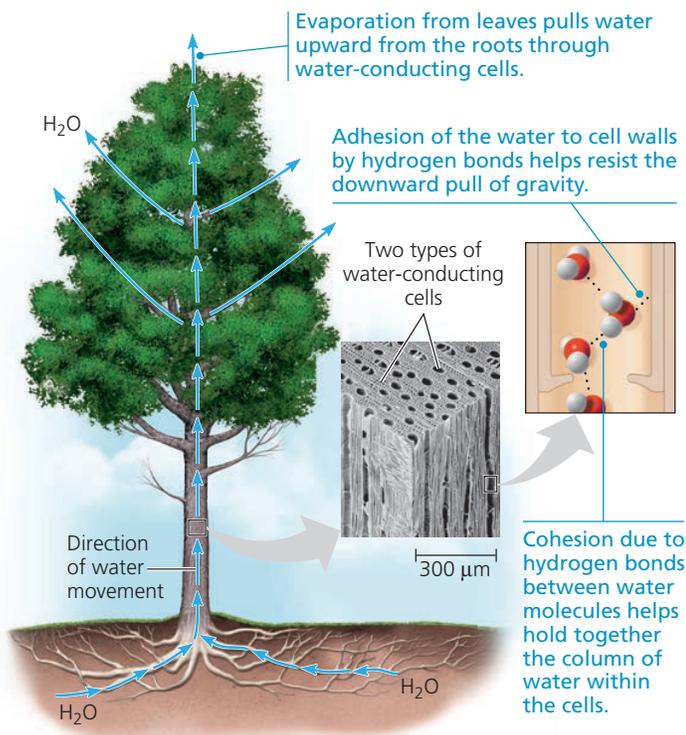
#### Cohesion of Water Molecules

Water molecules stay close to each other as a result of hydrogen bonding. Although the arrangement of molecules in a sample of liquid water is constantly changing, at any given moment many of the molecules are linked by multiple hydrogen bonds. These linkages make water more structured than most other liquids. Collectively, the hydrogen bonds hold the substance together, a phenomenon called **cohesion**.

Cohesion due to hydrogen bonding contributes to the transport of water and dissolved nutrients against gravity in plants. Water from the roots reaches the leaves through a network of water-conducting cells (Figure 3.3). As water evaporates from a leaf, hydrogen bonds cause water molecules leaving the veins to tug on molecules farther down, and the upward pull is transmitted through the water-conducting cells all the way to the roots. **Adhesion**, the clinging of one substance to another, also plays a role. Adhesion of water by hydrogen bonds to the molecules of cell walls helps counter the downward pull of gravity (see Figure 3.3).

Related to cohesion is **surface tension**, a measure of how difficult it is to stretch or break the surface of a liquid. At the interface between water and air is an ordered arrangement of water molecules, hydrogen-bonded to one another and to the water below, but not to the air above. This asymmetry gives

▼ **Figure 3.3 Water transport in plants.** Because of the properties of cohesion and adhesion, the tallest trees can transport water more than 100 m upward—approximately one-quarter the height of the Empire State Building in New York City.



**MB** BioFlix® Animation: Adhesion and Cohesion in Plants  
Animation: Cohesion of Water

water an unusually high surface tension, making it behave as though it were coated with an invisible film. You can observe the surface tension of water by slightly overfilling a drinking glass; the water will stand above the rim. The spider in **Figure 3.4** takes advantage of the surface tension of water to walk across a pond without breaking the surface.

## Moderation of Temperature by Water

Water moderates air temperature by absorbing heat from air that is warmer and releasing the stored heat to air that is cooler. Water is effective as a heat bank because it can absorb or release a relatively large amount of heat with only a slight change in its own temperature. To understand this capability of water, let's first look at temperature and heat.



◀ **Figure 3.4 Walking on water.** The high surface tension of water, resulting from the collective strength of its hydrogen bonds, allows this raft spider to walk on the surface of a pond.

## Temperature and Heat

Anything that moves has **kinetic energy**, the energy of motion. Atoms and molecules have kinetic energy because they are always moving, although not necessarily in any particular direction. The faster a molecule moves, the greater its kinetic energy. The kinetic energy associated with the random movement of atoms or molecules is called **thermal energy**. Thermal energy is related to temperature, but they are not the same thing. **Temperature** represents the *average* kinetic energy of the molecules in a body of matter, regardless of volume, whereas the thermal energy of a body of matter reflects the *total* kinetic energy, and thus depends on the matter's volume. When water is heated in a coffeemaker, the average speed of the molecules increases, and the thermometer records this as a rise in temperature of the liquid. The total amount of thermal energy also increases in this case. Note, however, that although the pot of coffee has a much higher temperature than, say, the water in a swimming pool, the swimming pool contains more thermal energy because of its much greater volume.

Whenever two objects of different temperature are brought together, thermal energy passes from the warmer to the cooler object until the two are the same temperature. Molecules in the cooler object speed up at the expense of the thermal energy of the warmer object. An ice cube cools a drink not by adding coldness to the liquid, but by absorbing thermal energy from the liquid as the ice itself melts. Thermal energy in transfer from one body of matter to another is defined as **heat**.

One convenient unit of heat used in this book is the **calorie (cal)**. A calorie is the amount of heat it takes to raise the temperature of 1 g of water by 1°C. Conversely, a calorie is also the amount of heat that 1 g of water releases when it cools by 1°C. A **kilocalorie (kcal)**, 1,000 cal, is the quantity of heat required to raise the temperature of 1 kilogram (kg) of water by 1°C. (The “Calories” on food packages are actually kilocalories.) Another energy unit used in this book is the **joule (J)**. One joule equals 0.239 cal; one calorie equals 4.184 J.

## Water's High Specific Heat

The ability of water to stabilize temperature stems from its relatively high specific heat. The **specific heat** of a substance is defined as the amount of heat that must be absorbed or lost for 1 g of that substance to change its temperature by 1°C. We already know water's specific heat because we have defined a calorie as the amount of heat that causes 1 g of water to change its temperature by 1°C. Therefore, the specific heat of water is 1 calorie per gram and per degree Celsius, abbreviated as 1 cal/(g · °C). Compared with most other substances, water has an unusually high specific heat. For example, ethyl alcohol, the type of alcohol in alcoholic beverages, has a specific heat of 0.6 cal/(g · °C); that is, only 0.6 cal is required to raise the temperature of 1 g of ethyl alcohol by 1°C.

Because of the high specific heat of water relative to other materials, water will change its temperature less than other

liquids when it absorbs or loses a given amount of heat. The reason you can burn your fingers by touching the side of an iron pot on the stove when the water in the pot is still lukewarm is that the specific heat of water is ten times greater than that of iron. In other words, the same amount of heat will raise the temperature of 1 g of the iron much faster than it will raise the temperature of 1 g of the water. Specific heat can be thought of as a measure of how well a substance resists changing its temperature when it absorbs or releases heat. Water resists changing its temperature; when it does change its temperature, it absorbs or loses a relatively large quantity of heat for each degree of change.

We can trace water's high specific heat, like many of its other properties, to hydrogen bonding. Heat must be absorbed in order to break hydrogen bonds; by the same token, heat is released when hydrogen bonds form. A calorie of heat causes a relatively small change in the temperature of water because much of the heat is used to disrupt hydrogen bonds before the water molecules can begin moving faster. And when the temperature of water drops slightly, many additional hydrogen bonds form, releasing a considerable amount of energy in the form of heat.

What is the relevance of water's high specific heat to life on Earth? A large body of water can absorb and store a huge amount of heat from the sun in the daytime and during summer while warming up only a few degrees. At night and during winter, the gradually cooling water can warm the air. This capability of water serves to moderate air temperatures in coastal areas (Figure 3.5). The high specific heat of water also tends to stabilize ocean temperatures, creating a favorable environment for marine life. Thus, because of its high specific heat, the water that covers most of Earth keeps temperature fluctuations on land and in water within limits that permit life. Also, because organisms are made primarily of water, they are better able to resist changes in their own temperature than if they were made of a liquid with a lower specific heat.

### Evaporative Cooling

Molecules of any liquid stay close together because they are attracted to one another. Molecules moving fast enough to

overcome these attractions can depart the liquid and enter the air as a gas (vapor). This transformation from a liquid to a gas is called vaporization, or evaporation. Recall that the speed of molecular movement varies and that temperature is the *average* kinetic energy of molecules. Even at low temperatures, the speediest molecules can escape into the air. Some evaporation occurs at any temperature; a glass of water at room temperature, for example, will eventually evaporate completely. If a liquid is heated, the average kinetic energy of molecules increases and the liquid evaporates more rapidly.

**Heat of vaporization** is the quantity of heat a liquid must absorb for 1 g of it to be converted from the liquid to the gaseous state. For the same reason that water has a high specific heat, it also has a high heat of vaporization relative to most other liquids. To evaporate 1 g of water at 25°C, about 580 cal of heat is needed—nearly double the amount needed to vaporize a gram of alcohol or ammonia. Water's high heat of vaporization is another emergent property resulting from the strength of its hydrogen bonds, which must be broken before the molecules can exit from the liquid in the form of water vapor.

The high amount of energy required to vaporize water has a wide range of effects. On a global scale, for example, it helps moderate Earth's climate. A considerable amount of solar heat absorbed by tropical seas is consumed during the evaporation of surface water. Then, as moist tropical air circulates poleward, it releases heat as it condenses and forms rain. On an organismal level, water's high heat of vaporization accounts for the severity of steam burns. These burns are caused by the heat energy released when steam condenses into liquid on the skin.

As a liquid evaporates, the surface of the liquid that remains behind cools down (its temperature decreases). This **evaporative cooling** occurs because the "hottest" molecules, those with the greatest kinetic energy, are the most likely to leave as gas. It is as if the 100 fastest runners at a college transferred to another school; the average speed of the remaining students would decline.

Evaporative cooling of water contributes to the stability of temperature in lakes and ponds and also provides a mechanism that prevents terrestrial organisms from overheating. For example, evaporation of water from the leaves of a plant helps keep the tissues in the leaves from becoming too warm in the sunlight. Evaporation of sweat from human skin dissipates body heat and helps prevent overheating on a hot day or when excess heat is generated by strenuous activity. High humidity on a hot day increases discomfort because the high concentration of water vapor in the air inhibits the evaporation of sweat from the body.

### Floating of Ice on Liquid Water

Water is one of the few substances that are less dense as a solid than as a liquid. In other words, ice floats on liquid water. While other materials contract and become denser when they solidify, water expands. The cause of this exotic behavior is, once again,

▼ **Figure 3.5** Temperatures for the Pacific Ocean and Southern California on an August day.



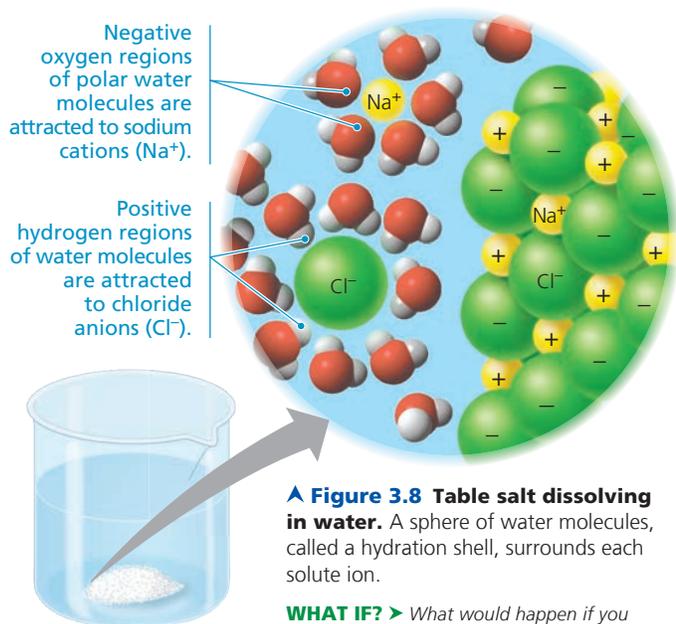
**INTERPRET THE DATA** ► Explain the pattern of temperatures shown in this diagram.

## Water: The Solvent of Life

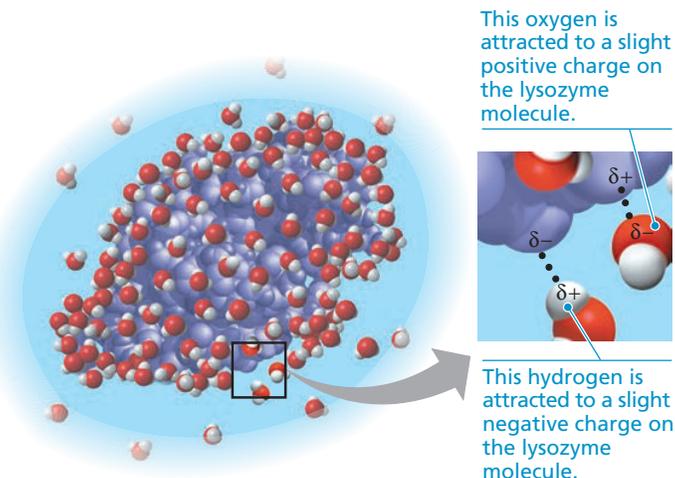
A sugar cube placed in a glass of water will dissolve with a little stirring. The glass will then contain a uniform mixture of sugar and water; the concentration of dissolved sugar will be the same everywhere in the mixture. A liquid that is a completely homogeneous mixture of two or more substances is called a **solution**. The dissolving agent of a solution is the **solvent**, and the substance that is dissolved is the **solute**. In this case, water is the solvent and sugar is the solute. An **aqueous solution** is one in which the solute is dissolved in water; water is the solvent.

Water is a very versatile solvent, a quality we can trace to the polarity of the water molecule. Suppose, for example, that a spoonful of table salt, the ionic compound sodium chloride (NaCl), is placed in water (**Figure 3.8**). At the surface of each crystal of salt, the sodium and chloride ions are exposed to the solvent. These ions and regions of the water molecules are attracted to each other due to their opposite charges. The oxygens of the water molecules have regions of partial negative charge that are attracted to sodium cations. The hydrogen regions are partially positively charged and are attracted to chloride anions. As a result, water molecules surround the individual sodium and chloride ions, separating and shielding them from one another. The sphere of water molecules around each dissolved ion is called a **hydration shell**. Working inward from the surface of each salt crystal, water eventually dissolves all the ions. The result is a solution of two solutes, sodium cations and chloride anions, homogeneously mixed with water, the solvent. Other ionic compounds also dissolve in water. Seawater, for instance, contains a great variety of dissolved ions, as do living cells.

A compound does not need to be ionic to dissolve in water; many compounds made up of nonionic polar molecules,



**Figure 3.9 A water-soluble protein.** Human lysozyme is a protein found in tears and saliva that has antibacterial action (see Figure 5.16). This model shows the lysozyme molecule (purple) in an aqueous environment. Ionic and polar regions on the protein's surface attract the partially charged regions on water molecules.



such as the sugar in the sugar cube mentioned earlier, are also water-soluble. Such compounds dissolve when water molecules surround each of the solute molecules, forming hydrogen bonds with them. Even molecules as large as proteins can dissolve in water if they have ionic and polar regions on their surface (**Figure 3.9**). Many different kinds of polar compounds are dissolved (along with ions) in the water of such biological fluids as blood, the sap of plants, and the liquid within all cells. Water is the solvent of life.

### Hydrophilic and Hydrophobic Substances

Any substance that has an affinity for water is said to be **hydrophilic** (from the Greek *hydro*, water, and *philos*, loving). In some cases, substances can be hydrophilic without actually dissolving. For example, some molecules in cells are so large that they do not dissolve. Another example of a hydrophilic substance that does not dissolve is cotton, a plant product. Cotton consists of giant molecules of cellulose, a compound with numerous regions of partial positive and partial negative charges that can form hydrogen bonds with water. Water adheres to the cellulose fibers. Thus, a cotton towel does a great job of drying the body, yet it does not dissolve in the washing machine. Cellulose is also present in the walls of water-conducting cells in a plant; you read earlier how the adhesion of water to these hydrophilic walls helps water move up the plant against gravity.

There are, of course, substances that do not have an affinity for water. Substances that are nonionic and nonpolar (or otherwise cannot form hydrogen bonds) actually seem to repel water; these substances are said to be **hydrophobic** (from the Greek *phobos*, fearing). An example from the kitchen is vegetable oil, which, as you know, does not mix stably with water-based substances such as vinegar. The hydrophobic

behavior of the oil molecules results from a prevalence of relatively nonpolar covalent bonds, in this case bonds between carbon and hydrogen, which share electrons almost equally. Hydrophobic molecules related to oils are major ingredients of cell membranes. (Imagine what would happen to a cell if its membrane dissolved!)

### Solute Concentration in Aqueous Solutions

Most of the chemical reactions in organisms involve solutes dissolved in water. To understand such reactions, we must know how many atoms and molecules are involved and calculate the concentration of solutes in an aqueous solution (the number of solute molecules in a volume of solution).

When carrying out experiments, we use mass to calculate the number of molecules. We must first calculate the **molecular mass**, which is the sum of the masses of all the atoms in a molecule. As an example, let's calculate the molecular mass of table sugar (sucrose),  $C_{12}H_{22}O_{11}$ , by multiplying the number of atoms by the atomic mass of each element (see Appendix B). In round numbers of daltons, the mass of a carbon atom is 12, the mass of a hydrogen atom is 1, and the mass of an oxygen atom is 16. Thus, sucrose has a molecular mass of  $(12 \times 12) + (22 \times 1) + (11 \times 16) = 342$  daltons. Because we can't weigh out small numbers of molecules, we usually measure substances in units called moles. Just as a dozen always means 12 objects, a **mole (mol)** represents an exact number of objects:  $6.02 \times 10^{23}$ , which is called Avogadro's number. Because of the way in which Avogadro's number and the unit *dalton* were originally defined, there are  $6.02 \times 10^{23}$  daltons in 1 g. Once we determine the molecular mass of a molecule such as sucrose, we can use the same number (342), but with the unit *gram*, to represent the mass of  $6.02 \times 10^{23}$  molecules of sucrose, or 1 mol of sucrose (this is sometimes called the *molar mass*). To obtain 1 mol of sucrose in the lab, therefore, we weigh out 342 g.

The practical advantage of measuring a quantity of chemicals in moles is that a mole of one substance has exactly the same number of molecules as a mole of any other substance. If the molecular mass of substance A is 342 daltons and that of substance B is 10 daltons, then 342 g of A will have the same number of molecules as 10 g of B. A mole of ethyl alcohol ( $C_2H_6O$ ) also contains  $6.02 \times 10^{23}$  molecules, but its mass is only 46 g because the mass of a molecule of ethyl alcohol is less than that of a molecule of sucrose. Measuring in moles makes it convenient for scientists working in the laboratory to combine substances in fixed ratios of molecules.

How would we make a liter (L) of solution consisting of 1 mol of sucrose dissolved in water? We would measure out 342 g of sucrose and then gradually add water, while stirring, until the sugar was completely dissolved. We would then add enough water to bring the total volume of the solution up to 1 L. At that point, we would have a 1-molar (1 M) solution of sucrose. **Molarity**—the number of moles of solute per liter of solution—is the unit of concentration most often used by biologists for aqueous solutions.

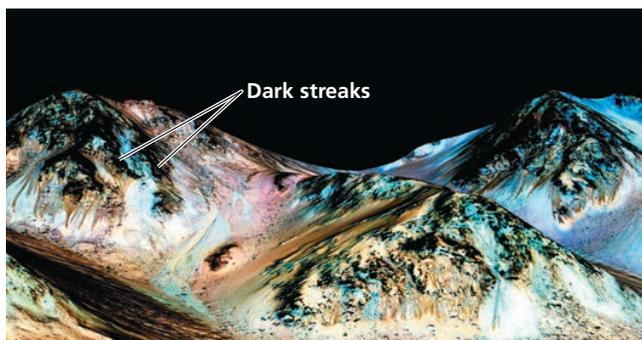
Water's capacity as a versatile solvent complements the other properties discussed in this chapter. Since these remarkable properties allow water to support life on Earth so well, scientists who seek life elsewhere in the universe look for water as a sign that a planet might sustain life.

 **MP3 Tutor: The Properties of Water**

### Possible Evolution of Life on Other Planets

**EVOLUTION** Biologists who look for life elsewhere in the universe (known as *astrobiologists*) have concentrated their search on planets that might have water. More than 800 planets have been found outside our solar system, and there is evidence for the presence of water vapor on a few of them. In our own solar system, Mars has been a focus of study. Like Earth, Mars has an ice cap at both poles. Images from spacecraft sent to Mars showed that ice is present just under the surface of Mars and enough water vapor exists in its atmosphere for frost to form. In 2015, scientists found evidence of water flowing on Mars (**Figure 3.10**), and other studies suggested conditions existed that could have supported microorganismal life. Drilling below the surface may be the next step in the search for signs of life on Mars. If any life-forms or fossils are found, their study will shed light on the process of evolution from an entirely new perspective.

**Figure 3.10 Evidence for liquid water on Mars.** Water appears to have helped form these dark streaks that run downhill on Mars during the summer. NASA scientists also found evidence of hydrated salts, indicating water is present. (This digitally treated photograph was taken by the Mars Reconnaissance Orbiter.)



### CONCEPT CHECK 3.2

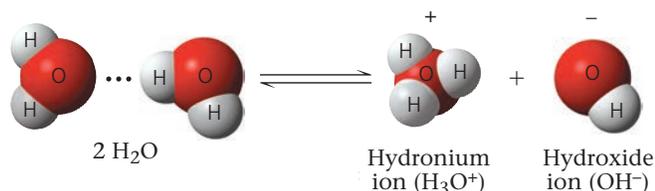
1. Describe how properties of water contribute to the upward movement of water in a tree.
2. Explain the saying "It's not the heat; it's the humidity."
3. How can the freezing of water crack boulders?
4. **WHAT IF? >** A water strider (an insect that can walk on water) has legs that are coated with a hydrophobic substance. What might be the benefit? What would happen if the substance were hydrophilic?
5. **INTERPRET THE DATA >** The concentration of the appetite-regulating hormone ghrelin is about  $1.3 \times 10^{-10}$  M in the blood of a fasting person. How many molecules of ghrelin are in 1 L of blood?

For selected answers, see Appendix A.

## CONCEPT 3.3

### Acidic and basic conditions affect living organisms

Occasionally, a hydrogen atom participating in a hydrogen bond between two water molecules shifts from one molecule to the other. When this happens, the hydrogen atom leaves its electron behind, and what is actually transferred is a **hydrogen ion** ( $H^+$ ), a single proton with a charge of  $1+$ . The water molecule that lost a proton is now a **hydroxide ion** ( $OH^-$ ), which has a charge of  $1-$ . The proton binds to the other water molecule, making that molecule a **hydronium ion** ( $H_3O^+$ ). We can picture the chemical reaction as follows:



#### Animation: Dissociation of Water Molecules

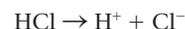
By convention,  $H^+$  (the hydrogen ion) is used to represent  $H_3O^+$  (the hydronium ion), and we follow that practice in this book. Keep in mind, though, that  $H^+$  does not exist on its own in an aqueous solution. It is always associated with a water molecule in the form of  $H_3O^+$ .

As indicated by the double arrows, this is a reversible reaction that reaches a state of dynamic equilibrium when water molecules dissociate at the same rate that they are being reformed from  $H^+$  and  $OH^-$ . At this equilibrium point, the concentration of water molecules greatly exceeds the concentrations of  $H^+$  and  $OH^-$ . In pure water, only one water molecule in every 554 million is dissociated; the concentration of  $H^+$  and of  $OH^-$  in pure water is therefore  $10^{-7} M$  (at  $25^\circ C$ ). This means there is only one ten-millionth of a mole of hydrogen ions per liter of pure water and an equal number of hydroxide ions. (Even so, this is a huge number—over 60,000 *trillion*—of each ion in a liter of pure water.)

Although the dissociation of water is reversible and statistically rare, it is exceedingly important in the chemistry of life.  $H^+$  and  $OH^-$  are very reactive. Changes in their concentrations can drastically affect a cell's proteins and other complex molecules. As we have seen, the concentrations of  $H^+$  and  $OH^-$  are equal in pure water, but adding certain kinds of solutes, called acids and bases, disrupts this balance. Biologists use something called the pH scale to describe how acidic or basic (the opposite of acidic) a solution is. In the remainder of this chapter, you will learn about acids, bases, and pH and why changes in pH can adversely affect organisms.

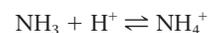
## Acids and Bases

What would cause an aqueous solution to have an imbalance in  $H^+$  and  $OH^-$  concentrations? When acids dissolve in water, they donate additional  $H^+$  to the solution. An **acid** is a substance that increases the hydrogen ion concentration of a solution. For example, when hydrochloric acid (HCl) is added to water, hydrogen ions dissociate from chloride ions:

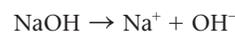


This source of  $H^+$  (dissociation of water is the other source) results in an acidic solution—one having more  $H^+$  than  $OH^-$ .

A substance that reduces the hydrogen ion concentration of a solution is called a **base**. Some bases reduce the  $H^+$  concentration directly by accepting hydrogen ions. Ammonia ( $NH_3$ ), for instance, acts as a base when the unshared electron pair in nitrogen's valence shell attracts a hydrogen ion from the solution, resulting in an ammonium ion ( $NH_4^+$ ):



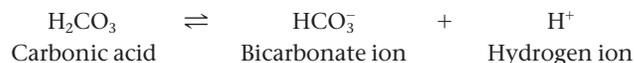
Other bases reduce the  $H^+$  concentration indirectly by dissociating to form hydroxide ions, which combine with hydrogen ions and form water. One such base is sodium hydroxide (NaOH), which in water dissociates into its ions:



In either case, the base reduces the  $H^+$  concentration. Solutions with a higher concentration of  $OH^-$  than  $H^+$  are known as basic solutions. A solution in which the  $H^+$  and  $OH^-$  concentrations are equal is said to be neutral.

Notice that single arrows were used in the reactions for HCl and NaOH. These compounds dissociate completely when mixed with water, so hydrochloric acid is called a strong acid and sodium hydroxide a strong base. In contrast, ammonia is a weak base. The double arrows in the reaction for ammonia indicate that the binding and release of hydrogen ions are reversible reactions, although at equilibrium there will be a fixed ratio of  $NH_4^+$  to  $NH_3$ .

Weak acids are acids that reversibly release and accept back hydrogen ions. An example is carbonic acid:



Here the equilibrium so favors the reaction in the left direction that when carbonic acid is added to pure water, only 1% of the molecules are dissociated at any particular time. Still, that is enough to shift the balance of  $H^+$  and  $OH^-$  from neutrality.

## The pH Scale

In any aqueous solution at  $25^\circ C$ , the product of the  $H^+$  and  $OH^-$  concentrations is constant at  $10^{-14}$ . This can be written

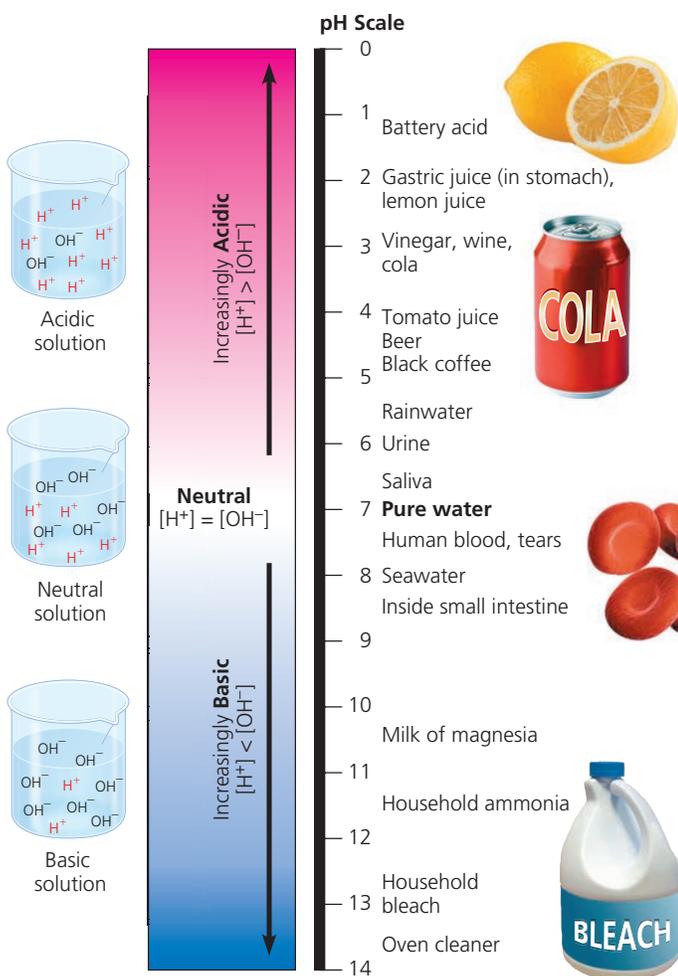
$$[H^+][OH^-] = 10^{-14}$$

(The brackets indicate molar concentration.) As previously mentioned, in a neutral solution at  $25^\circ C$ ,  $[H^+] = 10^{-7}$  and  $[OH^-] = 10^{-7}$ . Therefore, the product of  $[H^+]$  and  $[OH^-]$  in a

neutral solution at 25°C is  $10^{-14}$ . If enough acid is added to a solution to increase  $[H^+]$  to  $10^{-5} M$ , then  $[OH^-]$  will decline by an equivalent factor to  $10^{-9} M$  (note that  $10^{-5} \times 10^{-9} = 10^{-14}$ ). This constant relationship expresses the behavior of acids and bases in an aqueous solution. An acid not only adds hydrogen ions to a solution, but also removes hydroxide ions because of the tendency for  $H^+$  to combine with  $OH^-$ , forming water. A base has the opposite effect, increasing  $OH^-$  concentration but also reducing  $H^+$  concentration by the formation of water. If enough of a base is added to raise the  $OH^-$  concentration to  $10^{-4} M$ , it will cause the  $H^+$  concentration to drop to  $10^{-10} M$ . Whenever we know the concentration of either  $H^+$  or  $OH^-$  in an aqueous solution, we can deduce the concentration of the other ion.

Because the  $H^+$  and  $OH^-$  concentrations of solutions can vary by a factor of 100 trillion or more, scientists have developed a way to express this variation more conveniently than in moles per liter. The pH scale (Figure 3.11) compresses the range of  $H^+$  and  $OH^-$  concentrations by employing logarithms.

▼ **Figure 3.11** The pH scale and pH values of some aqueous solutions.



**Animation: Acids, Bases, and pH**

The **pH** of a solution is defined as the negative logarithm (base 10) of the hydrogen ion concentration:

$$\text{pH} = -\log [H^+]$$

For a neutral aqueous solution,  $[H^+]$  is  $10^{-7} M$ , giving us

$$-\log 10^{-7} = -(-7) = 7$$

Notice that pH *decreases* as  $H^+$  concentration *increases* (see Figure 3.11). Notice, too, that although the pH scale is based on  $H^+$  concentration, it also implies  $OH^-$  concentration. A solution of pH 10 has a hydrogen ion concentration of  $10^{-10} M$  and a hydroxide ion concentration of  $10^{-4} M$ .

The pH of a neutral aqueous solution at 25°C is 7, the midpoint of the pH scale. A pH value less than 7 denotes an acidic solution; the lower the number, the more acidic the solution. The pH for basic solutions is above 7. Most biological fluids, such as blood and saliva, are within the range of pH 6–8. There are a few exceptions, however, including the strongly acidic digestive juice of the human stomach (gastric juice), which has a pH of about 2.

Remember that each pH unit represents a tenfold difference in  $H^+$  and  $OH^-$  concentrations. It is this mathematical feature that makes the pH scale so compact. A solution of pH 3 is not twice as acidic as a solution of pH 6, but 1,000 times ( $10 \times 10 \times 10$ ) more acidic. When the pH of a solution changes slightly, the actual concentrations of  $H^+$  and  $OH^-$  in the solution change substantially.

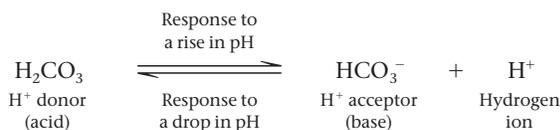
## Buffers

The internal pH of most living cells is close to 7. Even a slight change in pH can be harmful because the chemical processes of the cell are very sensitive to the concentrations of hydrogen and hydroxide ions. The pH of human blood is very close to 7.4, which is slightly basic. A person cannot survive for more than a few minutes if the blood pH drops to 7 or rises to 7.8, and a chemical system exists in the blood that maintains a stable pH. If 0.01 mol of a strong acid is added to a liter of pure water, the pH drops from 7.0 to 2.0. If the same amount of acid is added to a liter of blood, however, the pH decrease is only from 7.4 to 7.3. Why does the addition of acid have so much less of an effect on the pH of blood than it does on the pH of water?

The presence of substances called buffers allows biological fluids to maintain a relatively constant pH despite the addition of acids or bases. A **buffer** is a substance that minimizes changes in the concentrations of  $H^+$  and  $OH^-$  in a solution. It does so by accepting hydrogen ions from the solution when they are in excess and donating hydrogen ions to the solution when they have been depleted. Most buffer solutions contain a weak acid and its corresponding base, which combine reversibly with hydrogen ions.

Several buffers contribute to pH stability in human blood and many other biological solutions. One of these is carbonic

acid ( $\text{H}_2\text{CO}_3$ ), which is formed when  $\text{CO}_2$  reacts with water in blood plasma. As mentioned earlier, carbonic acid dissociates to yield a bicarbonate ion ( $\text{HCO}_3^-$ ) and a hydrogen ion ( $\text{H}^+$ ):



The chemical equilibrium between carbonic acid and bicarbonate acts as a pH regulator, the reaction shifting left or right as other processes in the solution add or remove hydrogen ions. If the  $\text{H}^+$  concentration in blood begins to fall (that is, if pH rises), the reaction proceeds to the right and more carbonic acid dissociates, replenishing hydrogen ions. But when the  $\text{H}^+$  concentration in blood begins to rise (when pH drops), the reaction proceeds to the left, with  $\text{HCO}_3^-$  (the base) removing the hydrogen ions from the solution and forming  $\text{H}_2\text{CO}_3$ . Thus, the carbonic acid–bicarbonate buffering system consists of an acid and a base in equilibrium with each other. Most other buffers are also acid–base pairs.

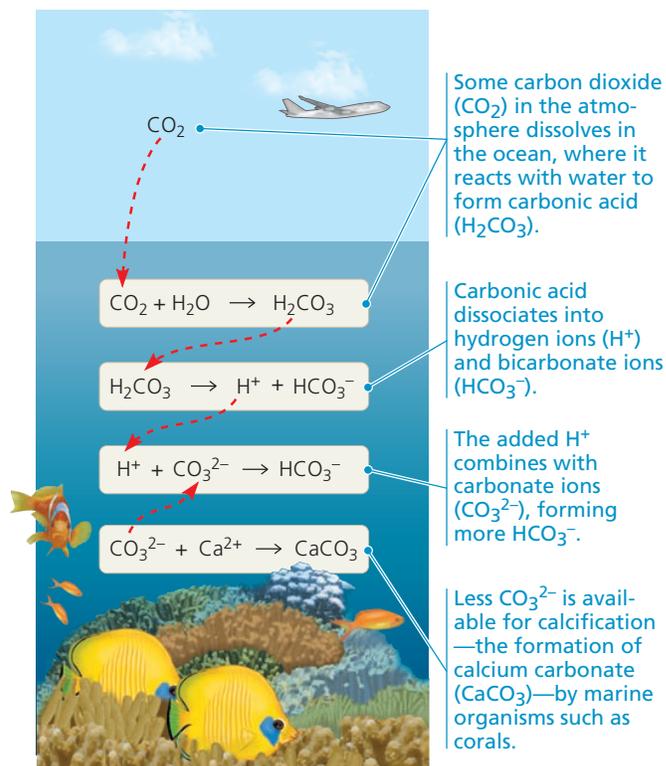
### Acidification: A Threat to Our Oceans

Among the many threats to water quality posed by human activities is the burning of fossil fuels, which releases  $\text{CO}_2$  into the atmosphere. The resulting increase in atmospheric  $\text{CO}_2$  levels has caused global warming and other aspects of climate change (see Concept 56.4). In addition, about 25% of human-generated  $\text{CO}_2$  is absorbed by the oceans. In spite of the huge volume of water in the oceans, scientists worry that the absorption of so much  $\text{CO}_2$  will harm marine ecosystems.

Recent data have shown that such fears are well founded. When  $\text{CO}_2$  dissolves in seawater, it reacts with water to form carbonic acid, which lowers ocean pH. This process, known as **ocean acidification**, alters the delicate balance of conditions for life in the oceans (Figure 3.12). Based on measurements of  $\text{CO}_2$  levels in air bubbles trapped in ice over thousands of years, scientists calculate that the pH of the oceans is 0.1 pH unit lower now than at any time in the past 420,000 years. Recent studies predict that it will drop another 0.3–0.5 pH unit by the end of this century.

As seawater acidifies, the extra hydrogen ions combine with carbonate ions ( $\text{CO}_3^{2-}$ ) to form bicarbonate ions ( $\text{HCO}_3^-$ ), thereby reducing the carbonate ion concentration (see Figure 3.12). Scientists predict that ocean acidification will cause the carbonate ion concentration to decrease by 40% by the year 2100. This is of great concern because carbonate ions are required for calcification, the production of calcium carbonate ( $\text{CaCO}_3$ ) by many marine organisms, including reef-building corals and animals that build shells. The **Scientific Skills Exercise** allows you to work with data from an experiment examining the effect of carbonate ion concentration on coral reefs. Coral reefs are sensitive ecosystems that act as havens for a great diversity of marine life.

▼ **Figure 3.12 Atmospheric  $\text{CO}_2$  from human activities and its fate in the ocean.**



**VISUAL SKILLS** ► Looking at all the chemical equations above, summarize the effect of adding excess  $\text{CO}_2$  to the oceans on the calcification process in the final equation.

**ABC News Video: Ocean Acidification**

The disappearance of coral reef ecosystems would be a tragic loss of biological diversity.

If there is any reason for optimism about the future quality of water resources on our planet, it is that we have made progress in learning about the delicate chemical balances in oceans, lakes, and rivers. Continued progress can come only from the actions of informed individuals, like yourselves, who are concerned about environmental quality. This requires understanding the crucial role that water plays in the suitability of the environment for continued life on Earth.

### CONCEPT CHECK 3.3

- Compared with a basic solution at pH 9, the same volume of an acidic solution at pH 4 has \_\_\_\_\_ times as many hydrogen ions ( $\text{H}^+$ ).
- $\text{HCl}$  is a strong acid that dissociates in water:  $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$ . What is the pH of 0.01 M  $\text{HCl}$ ?
- Acetic acid ( $\text{CH}_3\text{COOH}$ ) can be a buffer, similar to carbonic acid. Write the dissociation reaction, identifying the acid, base,  $\text{H}^+$  acceptor, and  $\text{H}^+$  donor.
- WHAT IF?** ► Given a liter of pure water and a liter solution of acetic acid, what would happen to the pH, in general, if you added 0.01 mol of a strong acid to each? Use the reaction from question 3 to explain the result.

For suggested answers, see Appendix A.